AP Chemistry Practice

Atomic Electron Configurations and Chemical Periodicity

Name: ____________________________

Date: __________________ Period: ____

1. Must all atoms having an odd atomic number be paramagnetic? Must all atoms having an even atomic number be diamagnetic? Explain.

2. Which of the following orbital diagrams are possible for a ground-state electron configuration and which are not? If the orbital diagram is not allowed state why it is not.

a) ↑↓ ↑↓ ↑↓ ↑↑ ↑↓
1s 2s 2p

b) ↑↓ ↑↓ ↑↓ ↑ ↑↓↑
1s 2s 2p

c) ↑↓ ↑↓ ↑ ↑ ↑
1s 2s 2p

d) ↑↓ ↑↓ ↓ ↓ ↓
1s 2s 2p

e) ↑↓ ↑↓ ↑ ↓ ↑
1s 2s 2p

3. What principle(s) or rule(s) does each of the following electron configurations violate?
a) $1s^2 2s^6 3s^2$
b) $1s^2 2s^2 2p^7 3s^1$
c) $1s^2 2s^2 2p^6 2d^3$

4. Use orbital box diagrams to represent the electron configurations of
a) Br^-
b) Ni^{2+}
c) Sb^{3+}
d) Te^{2-}
(Part of your diagram may be a noble gas core abbreviation)

5. Use the relationship between electron configurations and the periodic table to determine the number of:
a) valence shell electrons in an atom of Bi.
b) electrons in the fourth principal shell of Au
c) elements whose atoms have five valence shell electrons
d) unpaired electrons in an atom of Se
e) transition elements in the fifth period.

6. Which of the following are diamagnetic and which are paramagnetic? Explain.
a) an S atom
b) a Ba atom
c) a V^{2+} ion
d) an O^{2-} ion
e) an Ag atom

7. By only referring to a periodic table determine which member of the following pairs has the larger radius. Explain.
a) Cl or S
b) Cl^- or S^{2-}
c) Al or Mg
d) Mg^{2+} or F^-

8. Describe the trend in successive ionization energies as electrons are removed one at a time from an aluminum atom. Why is there a big jump between I_3 and I_4?

9. Two atoms have the electron configurations $1s^2 2s^2 2p^6$ and $1s^2 2s^2 2p^6 3s^1$. The first ionization energy of one is 2080 kJ/mol, and that of the other is 496 kJ/mol. Match each ionization energy with one of the given electron configurations. Justify your choice.

10. In a hydrogen atom, the 1s subshell is at an energy of -1.31×10^3 kJ/mol, whereas in a helium atom it is at -2.37×10^3 kJ/mol. Explain why the level is not the same in the two cases. For hydrogen-like atoms, Bohr's equation can be written as

$$E_n = -(2.18\times10^{-18}\text{J})Z^2(1/n^2)$$

(Where n is the principal quantum number and Z is the atomic number of the element). Calculate the energy of the 1s subshell for He^+ using this equation, and explain why the value you get is not the same as that given for helium.

11. Arrange the following isoelectronic species in order of a) increasing ionic radius and b) increasing ionization energy: O^{2-}, F^-, Na^+, Mg^{2+}

12. In general, as you move across the periodic table, the electron affinity of the elements becomes more negative. One exception to this trend, however, occurs when going from Group 4A elements to those in Group 5A. Refer to orbital box diagrams to explain why this exception is plausible and expected.

13. Silicon has an electron affinity of -134 kJ/mol. the electron affinity of phosphorus is -72 kJ/mol. Give a plausible reason for this difference.

14. Ionization energies are normally given in the unit of kJ/mol. Another way of expressing these quantities is in terms of a single atom rather than a mole of atoms, using the unit electron volts per atom, eV/atom. Use physical constants and other data from the appendices where necessary to show that 1eV/atom = 96.49 kJ/mol

15. Calculate I_1 for hydrogen
a) using Bohr's equation for the energy levels of the hydrogen atom, $E_n = -2.18\times10^{-18}\text{J}/n^2$, and
b) using the shortest wavelength line in the Balmer series of the hydrogen spectrum and the longest wavelength line in the Lyman series. Explain why they give the same result.